WJEC (Wales) Chemistry A-level

Topic 3.9 - Acid-Base Equilibria

Flashcards

What is the Lowry-Brønsted theory?

What is the Lowry-Brønsted theory?

The Lowry-Brønsted theory states that acid-base equilibria involves the transfer of protons between substances and substances can be classified as acids or bases depending on their interaction with protons.

Define a Lowry-Brønsted acid and give an example

Define a Lowry-Brønsted acid and give an example

A Lowry-Brønsted acid is a proton donor.

Example: Ammonium ions (NH_4^+)

Define a Lowry-Brønsted base and give an example

Define a Lowry-Brønsted base and give an example

A Lowry-Brønsted base is a proton acceptor.

Example: Hydroxide ions (OH^-)

Describe the difference between a strong acid and a weak acid

Describe the difference between a strong acid and a weak acid

A strong acid dissociates almost completely in water which means nearly all the H^+ ions are released.

A weak acid only partially dissociates in water so only a small number of H^+ ions are released.

Describe the difference between a strong base and a weak base

Describe the difference between a strong base and a weak base

A strong base dissociates almost completely in water so nearly all the OH^- ions are released.

A weak base only partially dissociates in water so only a small number of OH^- ions are released.

Give an expression for pH in terms of $[H^+]$

Give an expression for pH in terms of $[H^+]$

$$pH = -\log_{10}[H^+]$$

Equivalently:

$$[H^+] = 10^{-pH}$$

What is the relationship between pH and hydrogen ion concentration, $[H^+]$?

What is the relationship between pH and hydrogen ion concentration, $[H^+]$?

The pH scale is a measure of hydrogen ion concentration. The lower the pH, the higher the concentration of hydrogen ions.

What is the hydrogen ion concentration of a solution of hydrochloric acid which has a pH of 2.0?

What is the hydrogen ion concentration of a solution of hydrochloric acid which has a pH of 2.0?

$[H^+] = 10^{-pH}$

$= 10^{-2}$

$= 0.01 \text{ mol dm}^{-3}$

Give examples of strong acids and state the pH range which indicates a strong acid

Give examples of strong acids and state the pH range which indicates a strong acid

Examples:

Hydrochloric acid (HCl), sulfuric acid (H_2SO_4), nitric acid (HNO_3)

pH range of strong acids: 0-3

Give examples of weak acids and state the pH range which indicates a weak acid

Give examples of weak acids and state the pH range which indicates a weak acid

Examples:

Ethanoic acid (CH_3COOH), hydrogen sulfide (H_2S), any organic carboxylic acid

pH range of weak acids: 4 to just below 7

Give examples of strong bases and state the pH range which indicates a strong base

Give examples of strong bases and state the pH range which indicates a strong base

Examples:

Sodium hydroxide ($NaOH$), potassium hydroxide (KOH), calcium hydroxide ($Ca(OH)_2$)

pH range of strong bases: 12-14

Give examples of weak bases and state the pH range which indicates a weak base

Give examples of weak bases and state the pH range which indicates a weak base

Examples:

Ammonia (NH_3), methylamine (CH_3NH_2)

pH range for weak bases: just above 7 up to 11

What is the acid dissociation constant, Ka?

What is the acid dissociation constant, Ka?

The acid dissociation constant, Ka, is a measure of how strong an acid is in a solution.

Give the formula used to calculate Ka for a reaction of the form

$$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}$$

Give the formula used to calculate Ka for a reaction of the form $HA_{(aq)} \rightleftharpoons H^{+}_{(aq)} + A^{-}_{(aq)}$

$$Ka = \frac{[H^+][A^-]}{[HA]}$$

What are the units for Ka?

PMT
• resources • tuition • courses

What are the units for Ka?

mol dm^{-3}

How does the strength of an acid relate to the value of Ka?

How does the strength of an acid relate to the value of Ka?

Ka is the equilibrium constant for the dissociation of an acid: $HA \rightleftharpoons H^+ + A^-$

The stronger the acid, the further to the right the equilibrium lies so there is a higher concentration of products. This causes Ka to increase.

Why is Ka used to find the pH of a weak acid?

Why is Ka used to find the pH of a weak acid?

Weak acids only partially dissociate in water so the concentration of H^+ ions is not the same as the acid concentration (as with strong acids). This means the pH cannot be found using $[H^+]$ and so Ka is used instead.

Give the formula used to find Ka of a weak acid

Give the formula used to find Ka of a weak acid

For a weak acid you can assume that all the H^+ ions in solution come from the acid so that $[H^+_{(aq)}] \approx [A^-_{(aq)}]$ and you can assume that $[HA]_{equilibrium} \approx [HA]_{start}$

So, for a weak acid:

$$Ka = \frac{[H^+]^2}{[HA]}$$

Given that Ka for propanoic acid at 298 K is 1.3×10^{-5} mol dm^{-3}, what is the pH of 0.02 mol dm^{-3} of propanoic acid at 298 K?

Given that Ka for ethanoic acid at 298 K is 1.78 x 10^{-5} mol dm^{-3}, what is the pH of 0.02 mol dm^{-3} of ethanoic acid at 298 K?

First rearrange the equation for Ka to find $[H^+]$:

$$Ka = \frac{[H^+]^2}{[CH_3COOH]} \longrightarrow [H^+] = \sqrt{(Ka \times [CH_3COOH])}$$

So, $pH = -\log_{10}[H^+] = -\log_{10}(\sqrt{(1.78 \times 10^{-5} \times 0.02)}) = 3.22$

What is the difference between describing an acid/base as ‘concentrated’ compared to ‘strong’?

What is the difference between describing an acid/base as ‘concentrated’ compared to ‘strong’?

‘Concentrated’ implies there are many moles per dm^3.

‘Strong’ relates to the dissociation of the substance and implies the acid/base almost completely dissociates in water.

Give the formulas used to convert between Ka and pKa

Give the formulas used to convert between Ka and pKa

$$pKa = -\log_{10}(Ka)$$

$$Ka = 10^{-pKa}$$

What is the pKa of an acid which has a Ka value of 1.60×10^{-2} mol dm^{-3}?

What is the pKa of an acid which has a Ka value of 1.60×10^{-2} mol dm^{-3}?

$$pKa = -\log_{10}(Ka)$$

$$= -\log_{10}(1.6 \times 10^{-2})$$

$$= 1.80 \text{ (3.s.f)}$$

Give the equation for the ionic product of water

Give the equation for the ionic product of water

$$K_w = [H^+] [OH^-]$$

Or equivalently:

$$K_w = [H_3O^+] [OH^-]$$

Derive the ionic product of water using the equation for the ionisation of water

Derive the ionic product of water using the equation for the ionisation of water

In water, the following equilibrium is set up:

$$H_2O \rightleftharpoons H^+ + OH^-$$

So, $K_c = ([H^+][OH^-]) / [H_2O]$. Since $[H_2O]$ is very large compared to $[H^+]$ and $[OH^-]$, $[H_2O]K_c$ can be considered to be constant. Then $[H_2O]K_c = K_w$ and so $K_w = [H^+][OH^-]$.

What is the pH of pure water at room temperature?

What is the pH of pure water at room temperature?

7

How does the pH at which water is neutral ([OH] = [H+]) change as temperature increases?

How does the pH at which water is neutral ([OH] = [H+]) change as temperature increases?

In the equilibrium of water, the forwards reaction is endothermic and is therefore favoured when the temperature of the water is increased. So, the pH at which water is neutral (when the concentration of H^+ and OH^- is the same) decreases when the temperature is increased.

Given that K_w at 298 K is 1.0×10^{-14} mol^2 dm^{-6}, what is the pH of 0.10 mol dm^{-3} NaOH at 298 K?

Given that K_w at 298 K is 1.0×10^{-14} mol^2 dm^{-6}, what is the pH of 0.10 mol dm^{-3} NaOH at 298 K?

$K_w = [H^+] [OH^-] \rightarrow [H^+] = (K_w) / [OH^-]$

So, $[H^+] = (1.0 \times 10^{-4}) / 0.1 = 1.0 \times 10^{-13}$ mol dm^{-3}

Finally, $pH = -log_{10}[H^+] = -log_{10}(1 \times 10^{-13}) = 13.0$

What is a pH curve?

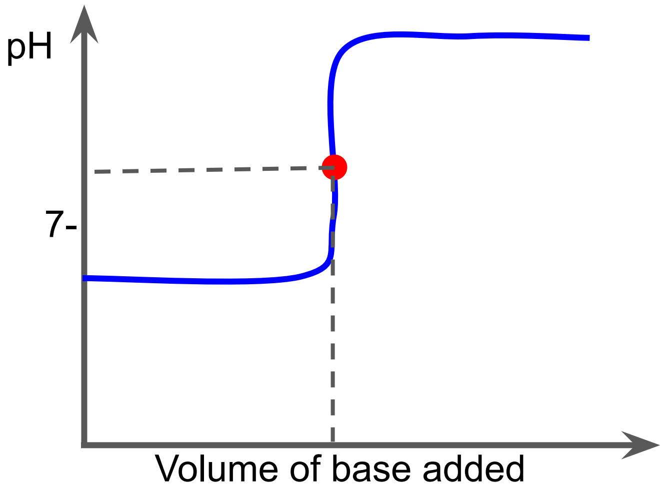
What is a pH curve?

Graphs which plot pH against volume of acid or base added are called pH curves and can be used to help identify the point of neutralisation of a solution.

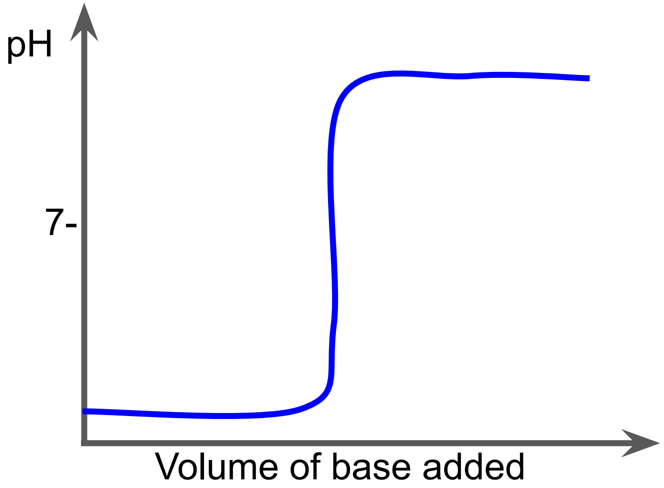
What is the equivalence point on a pH curve?

What is the equivalence point on a pH curve?

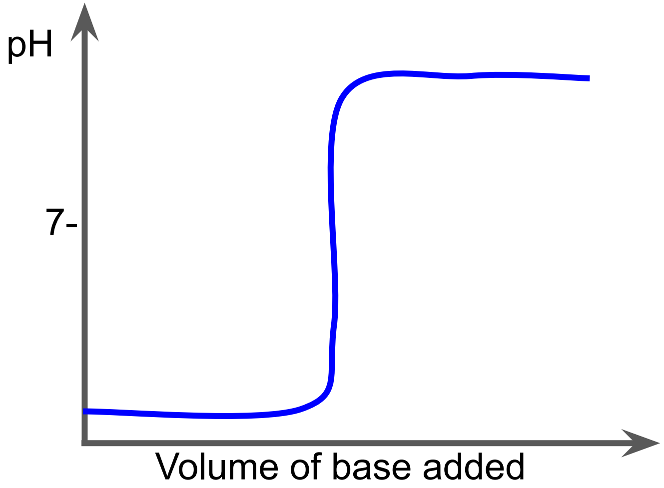
The equivalence point is also called the end point and is the point at which the pH curve is vertical. This is the point at which the solution has been neutralised.



Use the pH curve to identify the type of reactants involved in the reaction



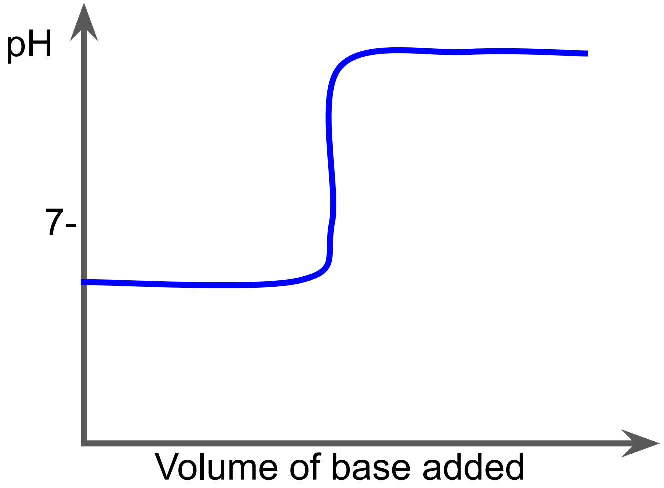
Use the pH curve to identify the type of reactants involved in the reaction



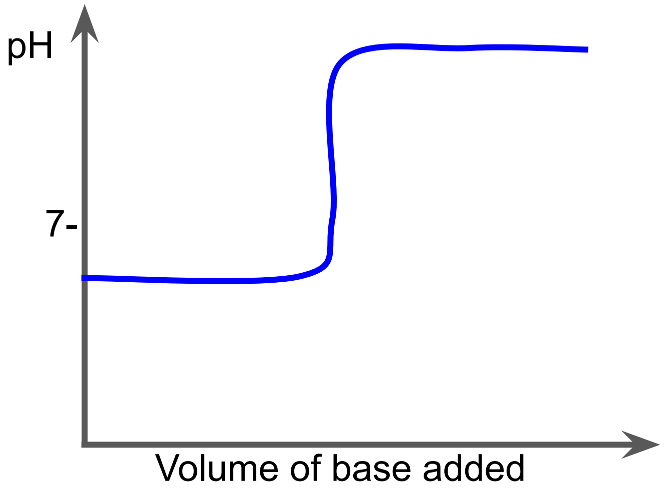
Strong base added to strong acid:

The pH is initially at around 1 as the strong acid is in excess. The pH ends up being very high, indicating a strong base is in excess.

Use the pH curve to identify the type of reactants involved in the reaction



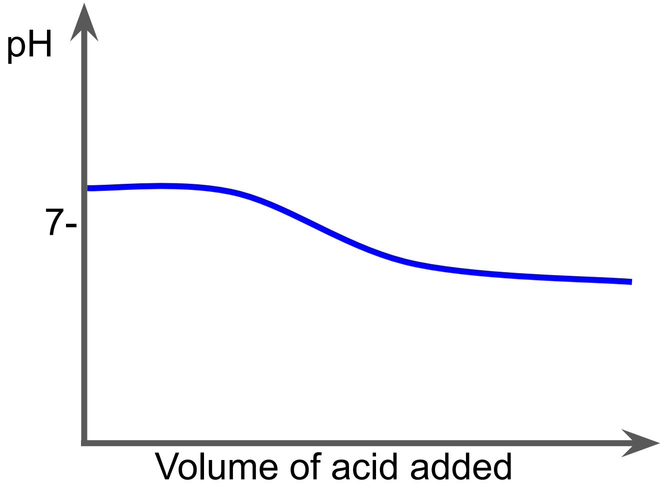
Use the pH curve to identify the type of reactants involved in the reaction



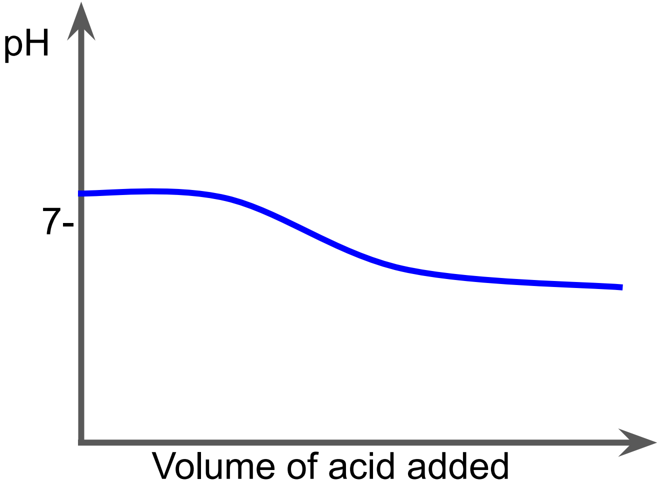
Strong base added to weak acid:

The pH starts about 5 where there’s an excess of weak acid. It finishes with a high pH where there’s an excess of strong base.

Use the pH curve to identify the type of reactants involved in the reaction



Use the pH curve to identify the type of reactants involved in the reaction



Weak acid added to weak base:

The pH starts about 8-9 where there's an excess of weak base. It finishes with a pH around 5 when there's an excess of weak acid.

What is a buffer solution?

What is a buffer solution?

A buffer solution is a solution which is able to resist changes in pH when small volumes of acid or base are added.

A buffer solution is commonly formed from a weak acid and its salt or a weak base and its salt. This produces a mixture containing H^+ ions and a large pool of OH^- ions which helps to resist any change in pH.

How is an acidic buffer solution made by mixing sodium ethanoate with ethanoic acid?

How is an acidic buffer solution made by mixing sodium ethanoate with ethanoic acid?

Ethanoic acid is a weak acid so will partially dissociate so lots of undissociated ethanoic acid molecules will remain in solution. The sodium ethanoate will fully dissociate, producing lots of ethanoate ions, CH_3COO^-. Therefore the following equilibrium is set up which will move to counteract changes in pH:

$$CH_3COOH_{(aq)} \rightleftharpoons H^+_{(aq)} + CH_3COO^-_{(aq)}$$

Consider the buffer solution made by mixing sodium ethanoate with ethanoic acid. How does the pH change when a small amount of acid is added?

Consider the buffer solution made by mixing sodium ethanoate with ethanoic acid. How does the pH change when a small amount of acid is added?

The following equilibrium is set up within the buffer solution:

$$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$$

If a small amount of acid is added, the concentration of H^+ increases. Most of the extra H^+ ions combine with CH_3COO^- ions to form CH_3COOH so the equilibrium shifts to the left. This reduces the concentration of H^+ to near to its original value so the pH does not change.

Explain the significance of buffer solutions in nature

Explain the significance of buffer solutions in nature

Buffer solutions are common in nature in order to keep systems regulated. Enzymes in living organisms often require an optimum pH and this can be maintained by a buffer solution.

Why are buffers used in industrial processes?

Why are buffers used in industrial processes?

Industrial processes use buffer solutions to maintain the optimum reaction conditions for large scale manufacturing.

Give an example of where buffer solutions are used in industrial processes

Give an example of where buffer solutions are used in industrial processes

Buffers are used in fermentation. Buffer solutions are added before fermentation begins to prevent the solution becoming too acidic.

They are also used in fabric dyeing processes and to perform chemical analysis.

Define salt hydrolysis

Define salt hydrolysis

A reaction where one of the ions from a salt reacts with water to form an acidic or basic solution.

What needs to be considered when selecting an indicator for a titration?

What needs to be considered when selecting an indicator for a titration?

The indicator chosen for a titration must change colour at exactly the end point of the titration. The indicator should change colour over a narrow pH range, which coincides with the vertical section of the pH curve.

Give an example of an indicator which can be used in a titration and describe its colour change

Give an example of an indicator which can be used in a titration and describe its colour change

Indicator	Colour at low pH	pH of colour change	Colour at high pH
Phenolphthalein	Colourless	8.3-10	Pink
Methyl orange	Red	3.1-4.4	Yellow

Why should a pH meter be used instead of an indicator in weak acid/weak base titrations?

Why should a pH meter be used instead of an indicator in weak acid/weak base titrations?

In weak acid/weak base titrations there is no sharp pH change so it is very difficult to successfully use an indicator to get an accurate point of neutralisation. A pH meter should be used instead since it will be more accurate and precise.